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- Q1. (a)** In theory, air is a solution since it is a homogeneous mixture of gases, mostly nitrogen, oxygen, and argon. However, some students may recognize that the presence of particulates such as soot, pollen, dust, and water vapour make a heterogeneous mixture.
- (b)** Solution
 - (c)** Pure substance
 - (d)** Solution
 - (e)** Element
 - (f)** Heterogeneous mixture
- Q2. (a)** Solvent: water; solute: carbonic acid
- (b)** Solvent: water; solute: ethylene glycol
 - (c)** Solvent: iron; solute: nickel
 - (d)** Solvent: nitrogen; solute: oxygen, argon

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- Q3.** Step 1: Bonds broken between molecules of solute, endothermic.
- Step 2: Bonds broken between molecules of solvent, endothermic.
- Step 3: Bonds formed between solute and solvent molecules, exothermic.
- Q4.** If the amount of energy released during the formation of the new bonds is greater than the energy required to break the bonds in steps 1 and 2, the process will be exothermic.
- Q5.** Students should mention that when the forces of attraction between solute molecules are weaker than those

between solute and solvent molecules, a substance dissolves. If the forces of attraction between solute molecules are stronger than those between solute and solvent molecules, the substance will not dissolve.

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- Q6. (a)** AgBr(s) (use solubility table)
(b) $\text{Mg(NO}_3)_2\text{(aq)}$ (because all nitrates are soluble)
(c) $\text{(NH}_4)_3\text{PO}_4\text{(aq)}$ (because all compounds with ammonium ions are soluble)
(d) $\text{CaCO}_3\text{(s)}$ (use solubility table for the rest)
(e) $\text{Ba(ClO}_3)_2\text{(aq)}$
(f) $\text{CsCH}_3\text{COO(aq)}$
(g) $\text{AlI}_3\text{(aq)}$
(h) $\text{KHCO}_3\text{(aq)}$

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- Q7.** Electrolytes: **(a)**, **(b)**, **(d)**, and **(h)**; non-electrolytes: **(c)** and **(f)**. Although $\text{Mg(OH)}_2\text{(s)}$ is only slightly soluble, a solution of $\text{Mg(OH)}_2\text{(s)}$ will conduct to a small degree.
- Q8.** When the reactants are dissolved in water, the solute particles separate, disperse, and collide with particles of other solutes, allowing the reaction to occur quickly.

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- 1. (a)** Methane is the solvent in natural gas, since it is the largest gaseous component (at 70–90%) of natural gas.
- (b)** Three solutes of natural gas are ethane, propane, and butane.

2. The moth pin is an example of a solution if it is made of more than one metal (an alloy). The pin made be made of brass (copper and zinc) or a solution of gold and other metals.
3. The sugar crystals contain sucrose molecules that require energy to separate. The heat energy from the hot coffee provides the energy for this endothermic process to occur more rapidly. Water molecules form bonds with the separated sucrose molecules causing the solution to release energy.
4. The amount of thermal energy released when the solvent (water) molecules bond with the solute (urea) molecules is much less than the amount of thermal energy used to break solvent-solvent bonds and solute-solute bonds.
5. In order for a chemical change to occur, particles must be separated and rearranged. Dissolving the particles in a solvent partly accomplishes this.
6. When hydrochloric acid is dissolved in water, it dissociates completely; most molecular compounds do not. Hydrochloric acid is a polar molecule, so there are dipole-dipole interactions between hydrochloric acid and water. The separation in charge creates a pathway for the flow of a current.
7. Dissolve a small amount of each of the solids in 50 mL of distilled water. Use a conductivity apparatus to test the solutions for conductivity. The sugar solution will not conduct because sugar is a molecular compound. The salt solution will conduct because salt is an ionic compound.

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- Q9. (a)** 188 g
(b) unsaturated
(c) 44 g
(d) 104 g

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Q10. (a) At equilibrium, the rate of the forward reaction or process must equal the rate of the reverse reaction or process in a closed system.

(b) In a saturated solution, the solid solute dissolves at the same rate as the aqueous solute precipitates.

Q11. (a) $\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$

(b) $\text{CaCl}_2(s) \rightarrow \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq)$

(c) $(\text{NH}_4)_2\text{CO}_3(s) \leftrightarrow 2\text{NH}_4^+(aq) + \text{CO}_3^{2-}(aq)$

(d) $\text{Mg}_3(\text{PO}_4)_2(s) \leftrightarrow 3\text{Mg}^{2+}(aq) + 2\text{PO}_4^{3-}(aq)$

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1. (a) Solubility is the amount of substance that dissolves in a given quantity of solvent at a given temperature to give a saturated solution.

(b) An unsaturated solution is a solution that is not in equilibrium and in which more of the solute can dissolve.

(c) A saturated solution is in equilibrium with respect to a dissolved substance, and no more solute can be dissolved at a given temperature.

2. (a) This solution would be considered unsaturated because the mass of solute, 20.0 g/100 mL, is lower than the given solubility.

(b) This is a supersaturated solution because it has a higher concentration than the solubility at this temperature.

3. (a) $\text{LiCl}(s) \leftrightarrow \text{Li}^+(aq) + \text{Cl}^-(aq)$

(b) $\text{NH}_4\text{F}(s) \leftrightarrow \text{NH}_4^+(aq) + \text{F}^-(aq)$

(c) $\text{Na}_2\text{S}(s) \leftrightarrow 2\text{Na}^+(aq) + \text{S}^{2-}(aq)$

(d) $(\text{NH}_4)_2\text{SO}_4(s) \leftrightarrow 2\text{NH}_4^+(aq) + \text{SO}_4^{2-}(aq)$

4. (a) Saturated solution
(b) $40.0\text{ g} - 34.7\text{ g} = 5.3\text{ g}$ undissolved solute
(c) $\text{KCl(s)} \leftrightarrow \text{K}^{\text{+}}(\text{aq}) + \text{Cl}^{-}(\text{aq})$
5. More solute can dissolve in the warmer thermal springs than in the cooler mountain springs.
6. The presence of the undissolved solute guarantees that the maximum amount of dissolved solute remains in solution as individual particles move from a dissolved to undissolved state at equilibrium.
7. (a) $\text{Ce}_2(\text{SO}_3)_3$ (b) $\text{Ce}_2(\text{SO}_3)_3$ (c) NaCl
(d) approximately $85\text{ }^{\circ}\text{C}$ (e) approximately 10.8 g

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1. (a) 20.7% (b) 16.61% (c) 15.1% 2. 35.5% 3. 85 g
4. 6.35 g 5. 15 g

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6. (a) 0.33 ppm (b) 3.3×10^2 ppb (c) $3.3 \times 10^{-5}\%$ (m/m)
7. 0.0700% (m/m) 8. 0.040 g 9. 1.4 ppm to 4.0 ppm
10. 0.0032 g

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11. (a) 1.7 mol/L (b) 2.41 mol/L
12. (a) 8.05×10^{-1} (b) 1.8×10^{-1}
13. 7.35×10^{-4} mol/L
14. 5.6×10^{-6} mol/L
15. 1.05×10^{-2}
16. 3.1 mol/L

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- 17. 0.29 mol/L
- 18. 3.2×10^{-7} mol/L
- 19. 2×10^{-6} mol/L
- 20. 3.45×10^{-7} mol/L
- 21. 2.4×10^{-2} mol/L
- 22. 1.00×10^{-2} mol/L

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- 23. (a) 1.46 g (b) 135 g
- 24. (a) 0.16 g (b) 1.5×10^2 g
- 25. 7.45 g
- 26. 4.5×10^2 g

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- 27. (a) 0.034 L (b) 0.052 L
- 28. (a) 0.0854 L (b) 3.1 L
- 29. 50 L
- 30. 1.0 L

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- 1. 50% *m/m*
- 2. (a) 27% *m/m*
(b) copper
(c) 73% of the sample is another element
- 3. 11 g
- 4. 36 ppb
- 5. $\text{ppb} = \mu\text{g/l} = 35 \mu\text{g}/0.500 \text{ L} = 70 \text{ ppb}$
No, I would not drink the water since the concentration of selenium is above the recommended limits.

6. 0.101 mol/L
7. $0.0024 \text{ mol/L Cu(NO}_3)_2(\text{aq}) \times \text{mol/L Cu(NO}_3)_2(\text{aq}) = 0.45 \text{ g Cu(NO}_3)_2 \times 1 \text{ mol/187.57 g} \times 1/0.100 \text{ L} = 0.0024 \text{ mol/L Cu(NO}_3)_2(\text{aq})$
8. $x \text{ mol/L KI (aq)} = 2.5 \text{ g KI(s)} \times 1 \text{ mol/166.00 g} \times 1/0.375 \text{ L} = 0.040 \text{ mol/L KI(aq)}$
9. $[\text{Mg}^{2+}(\text{aq})] = 3.42 \times 10^{-8} \text{ mol/L}$
 $[(\text{PO}_4)^-(\text{aq})] = 2.28 \times 10^{-8} \text{ mol/L}$
10. 3.73 g
11. 0.132 L
12. (a) ppm or ppb
 (b) mol/L
 (c) ppm or ppb
 (d) % *m/m*
 (e) % *m/m*
 (f) % *v/v*, % *m/m*
13. (a) Determine the mass of a dry beaker. Transfer the solution to the dry beaker and determine the mass of the beaker and solution. Subtract the mass of the beaker from the mass of the beaker and solution. This is the mass of the solution. Crystallize the solute dissolved in the solution

by either heating it slowly to evaporate the water, or allow the water to evaporate over a number of days. Mass the beaker with the dry solute. This is the mass of solute. Determine the percent by mass by dividing the mass of the dry solute by the mass of the solution and multiply by 100 to obtain a percentage.

- (b) To determine the molar concentration, you must know the molar mass. Either ask the teacher for the molar mass of the unknown compound, or ask the teacher for its chemical formula, and then calculate the molar mass. To determine molar concentration, use the mass of the solute divided by its molar mass to obtain the number of moles. Divide the number of moles by the volume of the solution. To obtain the volume of the solution, you can subtract the mass of the solute obtained in part a from the mass of the solution, also obtained in part a. This gives you the mass of the solvent, water. Since the density of water is 1.00 g/mL, convert the mass of the water to volume using density (1:1 ratio), and then divide by 1000 to get a volume in litres.
14. The concentration of a compound in solution is the number of moles of the compound per litre of solution. The concentration of an ion in solution depends on the concentration of the compound as well as the number of moles of the ion in the compound. The concentration of the potassium ion is three times that of the compound, so the formula is $K_3X(aq)$, where X could be $PO_4^{3-}(aq)$.
15. The technician must first decide what volume of solution to make. Since 5 groups of students will each be using 25 mL of the stock solution, a minimum of $5 \times 25 \text{ mL} = 125 \text{ mL}$ will be used. Since small amounts of solution will remain in measuring glassware, it is necessary to make a little extra. The technician might decide to make 150 mL of the stock solution. Next the technician must calculate the mass of solute that will result in a concentration of 1.5 mol/L.

First find the number of moles needed in 150 mL to make a concentration of 1.5 mol/L.

$$\begin{aligned}n &= cV \\&= \left(1.5 \frac{\text{mol}}{\text{L}}\right)(0.150 \text{ L}) \\&= 0.225 \text{ mol}\end{aligned}$$

Next, find the mass of sodium sulfate that contains 0.225 mol.

$$\begin{aligned}m &= nM \\&= (0.225 \text{ mol})\left(142.05 \frac{\text{g}}{\text{mol}}\right) \\&= 31.96 \text{ g}\end{aligned}$$

Use an electronic balance to tare the mass of a beaker and then add 32 g of sodium sulfate. Dissolve the sodium sulfate in approximately 60 mL of distilled water. Transfer the solution to a 150 mL volumetric flask using a funnel. Wash the beaker, stirring rod and funnel into the volumetric flask using distilled water. Mix the solution completely. Fill the volumetric flask with distilled water until the bottom of the

meniscus is at the line. Make sure to view the line at eye level and do not overfill. Stopper the flask and invert to mix.

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31. (a) 0.040 L **(b)** 0.128 L

32. (a) 0.300 mol/L **(b)** 0.0900 mol/L **(c)** 0.00720 mol/L

33. 0.021 mol/L

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1. A standard solution is one of precisely known concentration.

2. Step 1: Calculate the mass of solute required, based on a known volume and concentration of solution to be prepared.

Step 2: Using an electronic balance, obtain the required mass of solute in a beaker.

Step 3: Dissolve the solute in approximately half the distilled water volume required in the final solution. Stir until completely dissolved.

Step 4: Transfer the solution using a funnel into an appropriate-size volumetric flask.

Step 5: Rinse the funnel and beaker with distilled water and pour into the volumetric flask.

Step 6: Add distilled water to the flask up to the etched line.

Step 7: Cap and invert to mix.

3. Step 1: Calculate the volume of concentrated reagent required.

Step 2: Rinse the pipette with solution and discard in waste container.

Step 3: Transfer the required volume of concentrated solution into a volumetric flask.

Step 4: Fill the flask with distilled water to the etched line.

Step 5: Stopper the flask and invert to mix.

4. (a) The solution will be less concentrated because some of the solute was not added to the volumetric flask.
- (b) This solution will be more concentrated than desired because less water than expected is added to the solution.
- (c) Rinsing the pipette with water dilutes the standard solution that is pipetted to make the dilute solution. The final solution has a lower concentration than desired.

- (d) The solution will be more concentrated because of the additional solute added.
5. (a) Beaker, graduated cylinder, pipette. The markings on each piece of glassware dictate the accuracy of the volume measurement possible.
- (b) The pipette measures 10 mL of solution most precisely because it is the narrowest tube.
- (c) The beaker measures 10 mL of solution least precisely because it is the widest tube.
6. It is more accurate to measure a larger volume (that is, there is less relative error in measuring 10.0 mL than 1.0 mL or 0.1 mL)
7. 0.0125 mol/L
8. 1.5 L
- Transfer 350 mL of the 5.7 mol/L stock solution to a 1.5 L volumetric flask (either pouring the entire solution or transferring from a stock bottle using graduate or volumetric glassware). Rinse the transferring glassware into the volumetric flask. Fill the flask with distilled water until the bottom of the meniscus is on the etched line. Invert to mix.
9. 5.2 mL
- Approximately $150\text{ mL} - 5.3\text{ mL} = 144.7\text{ mL}$ would have to be added.
10. $(0.025\text{ L})(1.5\text{ mol/L}) = (0.050\text{ mol/L})V$
 $V = 0.75\text{ L}$
- Therefore, approximately $750\text{ mL} - 25\text{ mL} = 725\text{ mL}$ of water would have to be added.
11. $(475\text{ L})(5.5\text{ mol/L}) = (3.5 \times 10^{-7}\text{ mol/L})V$
 $V = 7.46 \times 10^9\text{ L}$

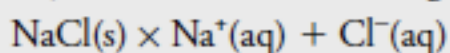
Answers to Understanding Concepts Questions

- (a) soda pop and juice

(b) hydrogen peroxide and mouthwash

(c) stainless steel sink and sterling silver jewelry
- A solution is a mixture of two or more substances; a pure substance is made up of only one type of element or compound.
- The turpentine is the solvent because it is present in the larger volume.
- Electrolytes ionize in water to form individual ions, whereas non-electrolytes remain as intact molecules.
- The dissolving process is exothermic when the energy released by the bonds forming between the solute and the solvent is greater than the energy required to separate the solid solutes and the solvent particles to make a solution. The opposite is true for the an endothermic solution.
- Insoluble suggests that a substance does not measurably dissolve at all. All substances dissolve to some extent in water.
- A saturated solution is in equilibrium with undissolved solute and contains the maximum amount of dissolved solute under the specific conditions. An unsaturated solution has less solute dissolved than the maximum possible. A supersaturated solution is unstable and contains a higher solute concentration than the solubility amount.
- A saturated solution contains both the solid and the dissolved form of the solute. A brine solution can be written as a system at equilibrium.

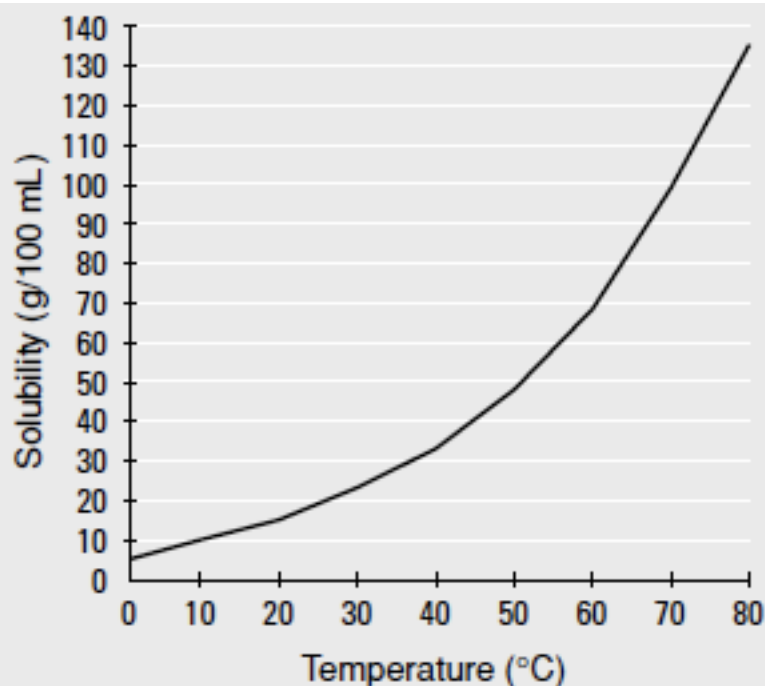
Crystallization $\leftrightarrow$ Dissolving



9. Liquid-in-liquid solutions are not affected appreciably by pressure and temperature. An increase in temperature increases the solubility of most solids but decreases the solubility of gases. An increase in pressure increases the solubility of gases in the solvent. The solubility of solids is not affected by pressure changes.

Answers to Applying Concepts Questions

10. The first test a student could perform is a conductivity test, to differentiate ionic and molecular solutions. Aqueous ionic compounds are electrolytes. Molecular compounds are non-electrolytes, and distilled water is also a non-electrolyte. The remaining two liquids can be heated to boil off the solvent. The one with no solid residue is the water and the other the molecular substance.
11. Controlled variables: Mass of solute, identity of solute, glassware used
Manipulated variable: Volume of water (concentration)
Responding variable: Temperature at which crystals appear
12. (a) The curve is shown below.



- (b) 86 g/100 mL (range of 82 to 88 g/100 mL is acceptable).
(c) 144 g/100 mL (range of 142–148 g/100 mL is acceptable).
(d) 77 °C (range of 75–78 °C is acceptable).
13. 400 g NaCl
14. 30 °C

Answers to Solving Problems Questions

15. 5.62 g
16. (a) 25 g (b) 225 g (c) 2.5 mol/L
17. 1.2 mg; 1.2 ppm
18. $[\text{Mg}^{2+}(\text{aq})] = 0.015 \text{ mol/L}$
19. (a) 0.172 mol/L
 (b) 0.577 mol/L
 (c) 3.03 mol/L
20. (a) $[\text{Mg}^{2+}(\text{aq})] = 1.53 \times 10^{-4} \text{ mol/L}$
 $[\text{OH}^{-}(\text{aq})] = 3.06 \times 10^{-4} \text{ mol/L}$
 (b) $[\text{Al}^{3+}(\text{aq})] = 7.4 \times 10^{-3} \text{ mol/L}$
 $[\text{NO}_3^{-}(\text{aq})] = 2.2 \times 10^{-2} \text{ mol/L}$
 (c) $[\text{PO}_4^{3-}(\text{aq})] = 1.83 \times 10^{-3} \text{ mol/L}$
 $[\text{Na}^{+}(\text{aq})] = 5.49 \times 10^{-3} \text{ mol/L}$
21. $[\text{NaCH}_3\text{COO}(\text{aq})] = 5.7 \text{ mol/L}$
22. (a) 7.89 g (b) 83 g
23. (a) 0.23 L (b) 10 L
24. (a) 1.7 mol/L (b) 1.44 mol/L
25. (a) 0.21 L (b) 4.20 mL

Answers to Making Connections Questions

26. The tea loses solubility with the decrease in temperature. Some of the tea has come out of solution as it cooled. The kinetic energy of the molecules has decreased.
27. True. Oil may sink or float on water depending on its density. Water is a polar substance and oil is non-polar. The two liquids do not form a homogeneous mixture.
28. Student speeches should include a number of perspectives on the issue. Scientific discussion could centre on the fact that sour gas contains sulfur, which when burned releases $\text{H}_2\text{S}(\text{g})$ into the air. This gas combines with water vapour to form acid rain. Increasing the acidity of rain water is problematic for living things. Environmental/ecological perspectives may indicate that sulfur gases have a rotten egg smell and may cause eye irritation and respiratory problems. The cost to scrub the sulfur emissions from the plant and to the health care system are economic considerations with potential legal implications. A proposal for why the plant should or should not be built can be based on a risk–benefit analysis of potential jobs and benefits to the community vs. the potential problems.
29. Activities may include: laundry, showers, toilet use, swimming, juice preparation, watering plants, washing the car. Volume estimates are difficult; however, many appliance companies have energy and/or volume rating information available. Two supplies of water may make environmental sense but are difficult in terms of practicality and economics. The current situation is likely due to economics, legal issues related to misuse, and the availability of technology to make two independent water systems viable.

- 30.** When the ammonium sulfate solution was sitting open, water was evaporating making the solution supersaturated. Because it was undisturbed and there were no existing crystals present to function as seeds for crystallization, no crystals formed. The motion caused by picking up the container disturbed the solution and caused crystallization to begin.